Assignment 17.2 Questions 10, 11, 35, 37, 55-59 odd

10) E^0 is the potential difference when under standard conditions (1.0M and 1.0 atm.) E is the potential difference when the cell is at any other concentration or partial pressure.

E is zero when the reaction reaches equilibrium (Q = K)

E^0 is zero whenever you have a concentration cell (same substance)

11) Anode: $Zn \longrightarrow Zn^{2+} + 2e^-$ $E^o = 0.76$ V

Cathode: $2x\,[Ag^+ + e^- \longrightarrow Ag]$ $E^o = 0.80$ V

$Zn(s) + 2Ag^+(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s)$ $E^o = 1.56$ V

$$Q = \frac{[Zn^{2+}]}{[Ag^+]^2} \qquad E = E^o - \frac{0.0592}{n} \log Q$$

As $[Zn^{2+}]$ increases, Q increases. If Q increases, E decreases (a little).
As $[Ag^+]$ increases, Q decreases. If Q decreases, E increases (a little).

35) a. Anode: $[\,2I^- \longrightarrow I_2 + 2\,e^-\,] \times 5$ $-E^o = -0.54$ V

Cathode: $[\,MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O\,]$ **x 2** $E^o = 1.51$ V

$2MnO_4^- + 16H^+ + 10I^- \longrightarrow 2Mn^{2+} + 8H_2O + 5I_2$ **E^o_{cell} = 0.97 V**

This is the balanced redox reaction, but you really don't need to do all of this. They are only showing you the species that are oxidized or reduced. You should be able to find the reduction potentials just from what they give you.

Because E^o_{cell} is positive, the reaction is spontaneous.

b. Anode: $2F^- \longrightarrow F_2 + 2\,e^-$ $-E^o = -2.87$ V

Cathode: $MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ $E^o = 1.51$ V

E^o_{cell} = -1.36 V

Because E^o_{cell} is negative, the reaction is non-spontaneous.

37. Na^+ is not gaining or losing electrons, so it is a spectator ion:

Anode: $2ClO_2^- \longrightarrow 2ClO_2 + 2e^-$ $-E^o = -0.954$ V

Cathode: $Cl_2 + 2e^- \longrightarrow 2Cl^-$ $E^o = 1.36$ V

E^o_{cell} = 0.41 V

$\Delta G^o = -nFE^o_{cell} = -\,(2 \text{ mol } e^-)(96{,}485)(\,0.406) =$ **- 78 kJ**

55) $E = E^o - \frac{0.0592}{n} \log Q$ $\quad Q = \frac{1}{[H^+]^2[HSO_4^-]^2} = \frac{1}{[4.5]^2[4.5]^2} = 0.00244$

$E = 2.04 - \frac{0.0592}{2} \log 0.00244 = 2.11\ V$

57)

Anode:	$Zn \longrightarrow Zn^{2+} + 2e^-$	$-E^o = 0.76\ V$
Cathode:	$Cu^{2+} + 2e^- \longrightarrow Cu$	$E^o = 0.34V$
	$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$	$E^o = 1.10\ V$

$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{[1.20]}{[0.80]} = 1.5$ $\quad E = 1.10 - \frac{0.0592}{2} \log 1.5 = 1.09\ V$

59) a.

Anode:	$H_2 \longrightarrow 2H^+ + 2e^-$	$-E^o = 0.00\ V$
Cathode:	$Cu^{2+} + 2e^- \longrightarrow Cu$	$E^o = 0.34V$
	$H_2(g) + Cu^{2+}(aq) \longrightarrow 2H^+(aq) + Cu(s)$	$E^o = 0.34\ V$

$Q = \frac{[H^+]^2}{[Cu^{2+}]} = \frac{[1.00]^2}{[0.00025]} = 4000$ $\quad E = 0.34 - \frac{0.0592}{2} \log 4000 = \mathbf{0.23\ V}$

b. $E = 0.34 - \frac{0.0592}{2} \log \frac{[1.00]^2}{[x]} = 0.195$

$-0.0296 \log \frac{1}{[x]} = -0.145$

$\log \frac{1}{[x]} = 4.90$

$\frac{1}{[x]} = 79{,}200$ $\quad \mathbf{[x] = 1.3 \times 10^{-5}\ M}$